TOPIC 09 – REDOX

9.2 – REDOX EQUATIONS

IB Chemistry

T09D01

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9.2 – Redox Equations

- 9.2.1 Deduce simple oxidation and reduction half-equations given the species involved in a redox reaction. (3)
  9.2.2 Deduce redox equations using half equations. (3)
  9.2.3 Define the terms oxidizing agent and reducing agent. (1)
  9.2.4 Identify the oxidizing and reducing agents in redox equations. (2)

Redox Half-Equations:

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9.2.1 Deduce simple oxidation and reduction half-equations given the species involved in a redox reaction. (3)

- Many redox reactions are only brought about under certain conditions.
- As discussed in Topic 10, an acidified aqueous solution can often provide these conditions
- The half equations frequently include H_2O and H^+ ions
- Similar procedures can be followed for the construction of half-equations for
  - Oxidation Reactions
  - Reduction Reactions

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For Oxidation Reactions

1. **Write down the formulae** of the reactant and products:
   - $Cr_2O_7^{2-} \rightarrow Cr^{3+}$
2. **Balance** with respect to the chromium:
   - $Cr_2O_7^{2-} \rightarrow \mathbf{2}Cr^{3+}$
3. **Balance the O** atoms of the dichromate (VI) **with H_2O**
   - $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$
4. **Balance the H** atoms in water with **H^+ ions**
   - $Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$
5. **Determine total charge** of each side
   - LHS: +14 -2 = +12 RHS: 2(+3) = +6
6. **Balance two charges by adding electrons**
   - $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
   - LHS: +14 -2 -6 = +6 RHS: (2x +3) = +6

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For Reduction Reactions

1. **Write down the formulae** of the reactant and products:
   - $HNO_2 \rightarrow NO_3^-$
2. **Balance** with respect to the nitrogen:
   - $HNO_2 \rightarrow NO_3^-$ (not needed)
3. **Balance the O** atoms of the dichromate (VI) **with H_2O**
   - HNO_2 **+ H_2O** $\rightarrow NO_3^-$
4. **Balance the H** atoms in water with **H^+ ions**
   - $HNO_2 + H_2O \rightarrow NO_3^-$ **+ $3H^+$**
5. **Determine total charge** of each side
   - LHS: 0 RHS: -1 + 3(+1) = -2
6. **Balance two charges by adding electrons**
   - $HNO_2 + H_2O \rightarrow NO_3^- + 3H^+$ **+ $2e^-$**
   - LHS: 0 RHS: -1 + 3(+1) -2 = 0

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Redox Equations:

9.2.2 Deduce redox equations using half equations. (3)

- If given both the reduction and oxidation half equations, you should be able to balance and simplify the equations
- A common example is the oxidation of a primary or secondary alcohol (via dichromate or manganate)

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If given both the half-equations

- Following Topic 10 again, if the half-equations for the oxidation of methanol (and reduction of manganate)
  - $2H_2O + CH_3OH \rightarrow CO_2 + H_2O + 6H^+ + 6e^-$
  - $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
- **Equalize the number of electrons** lost and gained
  - **5x (**$2H_2O + CH_3OH \rightarrow CO_2 + H_2O + 6H^+ + 6e^-$**)**
  - **6x (**$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$**)**
- **Add the two** half-equations together

$$\mathbf{10}H_2O + \mathbf{5}CH_3OH + \mathbf{6}MnO_4^- + \mathbf{48}H^+ + \mathbf{30}e^- \rightarrow \mathbf{5}CO_2 + \mathbf{5}H_2O + \mathbf{30}H^+ + \mathbf{30}e^- + \mathbf{6}Mn^{2+} + \mathbf{24}H_2O$$

- **Cancel electrons** and simplify numbers of H_2O and H^+
  - $5CH_3OH + 6MnO_4^- + \mathbf{18}H^+ \rightarrow 5CO_2 + 6Mn^{2+} + \mathbf{19}H_2O$

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Agents

9.2.3 Define the terms oxidizing agent and reducing agent. (1)

- An **oxidizing agent** is defined as a substance that brings about the oxidation of substances by accepting electrons from the substance they oxidize
  - Oxidizing agents undergo **reduction**
- A **reducing agent** is defined as a substance that brings about the reduction of a substance by donating electrons to the substance it reduces.
  - Reducing agents undergo **oxidation**

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Identify Agents

9.2.4 Identify the oxidizing and reducing agents in redox equations. (2)

- Common **reducing agents** (things that are **oxidized**)
  - Hydrogen: $O^{2-} + H_2 \rightarrow H_2O + \mathbf{2e^-}$
    - $PbO + H_2 \rightarrow H_2O + Pb$
  - Carbon: $O^{2-} + C \rightarrow CO + \mathbf{2e^-}$
    - $PbO + C \rightarrow Pb + CO$
  - Carbon Monoxide: $O^{2-} + CO \rightarrow CO_2 + \mathbf{2e^-}$
    - $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
  - Metals: $Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu$
    - $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$

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Oxidizing Agents

- Common **oxidizing agents** (things that are **reduced**)
  - Oxygen: ½ O_2 + **2e$^-$** → O^{2-}
  - Ozone: $O_3 + 2H^+ +$ **2e$^-$** → $O_2 + H_2O$
    - $2KI + O_3 + H_2O \rightarrow KOH + O_2 + I_2$
  - Chlorine: Cl_2 + **2e$^-$** → $2Cl^-$
  - Acidified $KMnO_4$: $MnO_4^- + 8H^+ +$ **5e$^-$** → $Mn^{2+} + 4H_2O$
  - Acidified $K_2Cr_2O_7$: $Cr_2O_7^{2-} + 14H^+ +$ **6e$^-$** → $2Cr^{3+} + 7H_2O$
  - Acidified H_2O_2: $H_2O_2 + 2H^+ +$ **2e$^-$** → $2H_2O$
  - Metal ions: $Ag^+ +$ **1e$^-$** → Ag^+
  - Hydrogen ions: $2H^+ +$ **2e$^-$** → H_2
    - $Mg + 2HCl \rightarrow MgCl_2 + H_2$
  - MnO_2: $MnO_2 + 4H^+ +$ **2e$^-$** → $Mn^{2+} + 2H_2O$

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Balancing via Redox Method

- Often, chemical equations are difficult to balance based due to large and uneven numbers of reacting and produced species.
- A systematic approach using the principles of redox may allow you to balance such equations:

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Balancing Redox Equations 1

1. Assign oxidation numbers to the species in the reaction
2. Find the substance oxidized and the substance reduced
3. Write half reactions for the oxidation and reduction
4. Balance the atoms that change in the half reaction
5. Determine the electrons transferred and balance the electrons between the half reactions
6. Combine the half reactions and balance the remaining atoms
7. Check your work. Make sure that both the atoms and charges balance

$\overset{0}{Cu} + \overset{+1\ \ +5\ \ -2}{HNO_3} \rightarrow \overset{+2\ \ +5\ \ -2}{Cu(NO_3)_2} + \overset{+2\ \ -2}{NO} + \overset{+1\ \ -2}{H_2O}$

Copper Oxidized, some Nitrogen Reduced

$Cu \rightarrow Cu^{2+} + 2e^-$ $\qquad$ $N^{5+} + 3e^- \rightarrow N^{2+}$

*both are balanced

2 e- ox, 3 e- red, lowest common multiple = 6 e-

$3Cu \rightarrow 3Cu^{2+} + 6e^-$ $\qquad$ $2N^{5+} + 6e^- \rightarrow 2N^{2+}$

$3Cu + 2N^{5+} + 6e^- \rightarrow 3Cu^{2+} + 6e^- + 2N^{2+}$

*Ratio of Cu compounds to NO (5+) must be 3:2

$3Cu + \_\_HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + \_\_H_2O$

*Next balance out the nitrogen (total)

$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + \_\_H_2O$

Balance the final product

$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

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Balancing Redox Equations 2

1. Assign oxidation numbers to the species in the reaction
2. Find the substance oxidized and the substance reduced
3. Write half reactions for the oxidation and reduction
4. Balance the atoms that change in the half reaction
5. Determine the electrons transferred and balance the electrons between the half reactions
6. Combine the half reactions and balance the remaining atoms
7. Check your work. Make sure that both the atoms and charges balance

+1 +5 -2 0 +1 +5 -2 +4 -2 +1 -2

$$HNO_3 + I_2 \rightarrow HIO_3 + NO_2 + H_2O$$

Iodine Oxidized, some Nitrogen Reduced

$I^0 \rightarrow I^{5+} + 5e^-$ $N^{5+} + 1e^- \rightarrow N^{4+}$

$I_2^0 \rightarrow 2I^{5+} + 10e^-$ $N^{5+} + 1e^- \rightarrow N^{4+}$

*both are balanced

10 e- ox, 1 e- red, lowest common multiple = 10e

$I_2^0 \rightarrow 2I^{5+} + 10e^-$ $10\ N^{5+} + 10e^- \rightarrow 10N^{4+}$

$$I_2^0 + 10\ N^{5+} + 10e^- \rightarrow 2I^{5+} + 10e^- + 10N^{4+}$$

*Ratio of N compounds to I must be 5:1

$5\ HNO_3 + 1\ I_2 \rightarrow 1\ HIO_3 + 5\ NO_2 + \_\_\ H_2O$

Next balance out the water, you're done

$$5\ HNO_3 + 1\ I_2 \rightarrow 1\ HIO_3 + 5\ NO_2 + 2\ H_2O$$

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Redox Titrations